Periodic Table Trends

Pre-AP

Target Words

- Valence Electron
- Alkali Metals
- Alkali Earth Metals
- Halogens
- Transition Metals
- Lanthanides
- Actinides
- Noble Gases
- Metal
- Nonmetal
- Metalloids
- Group
- Period
- Periodic Law
- Chemical/Physical Properties and Changes

Periodic Table History

- Atoms; Who was it that proposed the first scientific atomic theory?

- John Dalton – Where did all his empirical evidence come from?

- Atomic Weights – Relative weights in most cases.

Periodic Table History

- Atomic Weight was a big deal, several scientists made their names by compiling lists of weights.
- By the mid 1800's 70 or so elements had all been discovered and weighed, with some disagreement.
- Unfortunately, part of the problem was the several systems of notation and weights being used.

Karlsruhe Congress

- In 1860 some of the most prominent chemists in Europe called together the first international chemistry conference in Karlsruhe, Germany.
- Issues involving nomenclature, notation, and other topics were considered highly important.
- For example, acetic acid was represented by no less than 19 different proposed formulas!

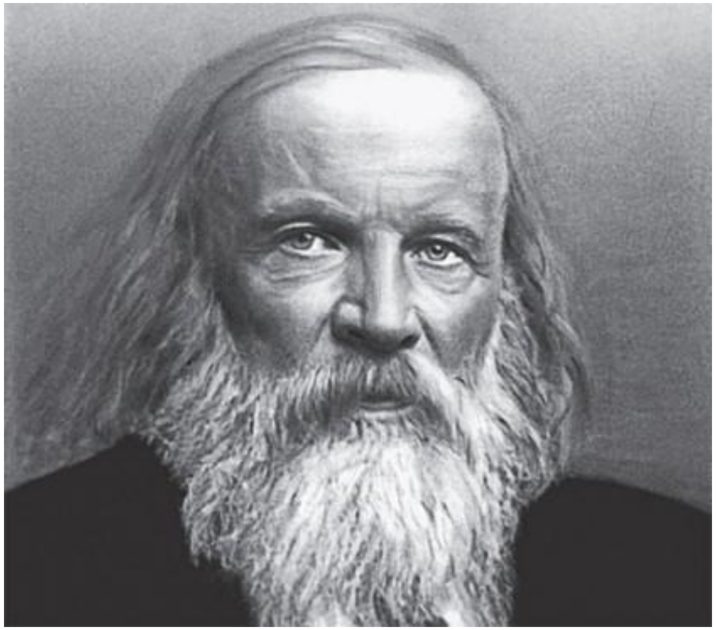
Karlsruhe Congress – Weighty Matters

- Perhaps the most important item was working on standardized atomic weights.

- Thanks have to go to Avagadro for some really important work and papers that he had released.

- By the end of the Congress, values for important elements were adopted as standard.
  - Hydrogen – 1
  - Carbon – 12
  - Oxygen – 16

Enter Mendeleev



- Russian Chemist, famously wrote *Principles of Chemistry* and in 1869 presented to the Russian Chemical Society his first draft of his extended **periodic table.**

- His table not only arranged the elements known at the time, but also predicted elements not yet discovered!

Mendeleev's Table

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Mendeleev's Table

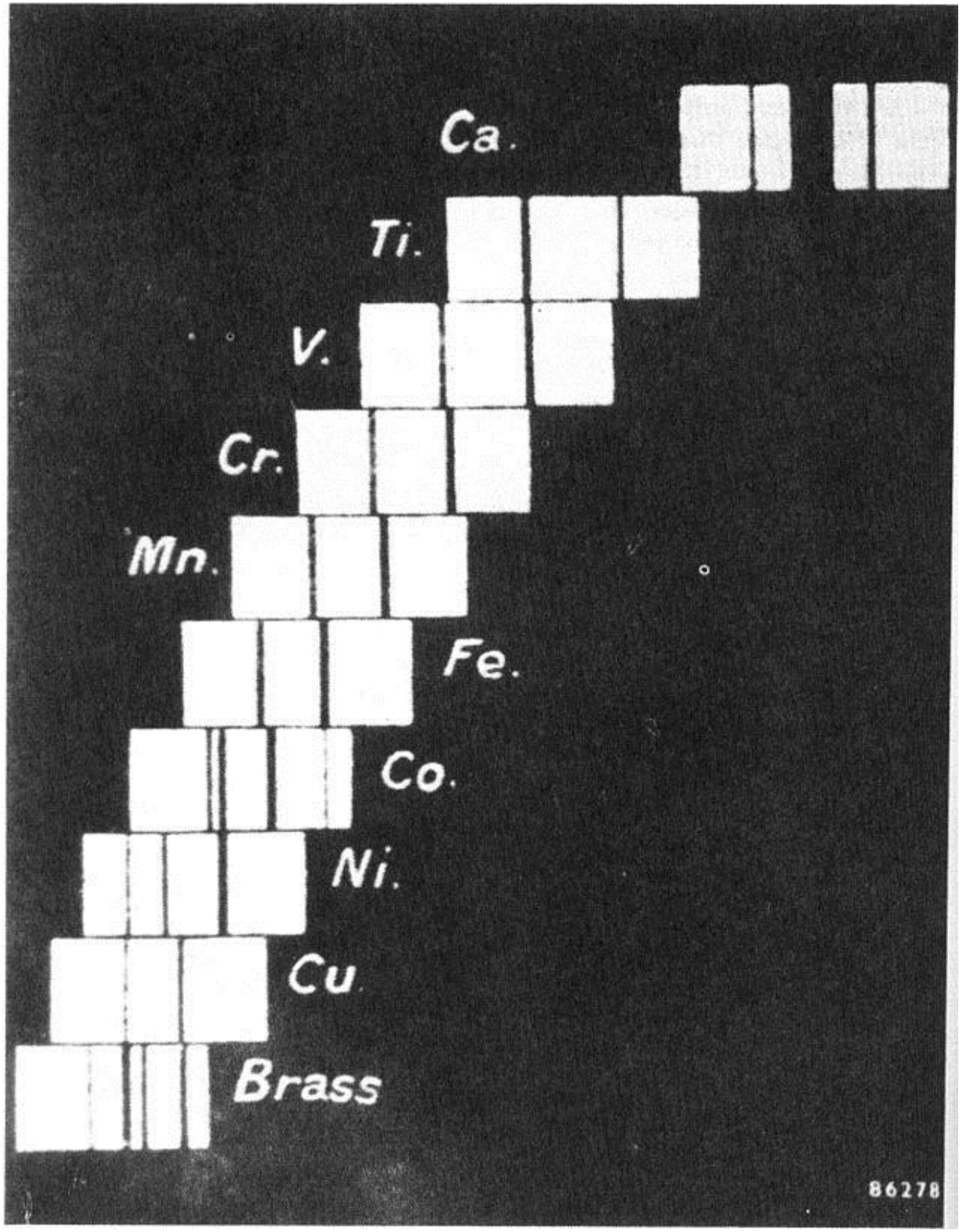
- Elements were placed in vertical columns according to **atomic mass numbers**
- Similar properties and characteristics seemed to repeat in certain patterns.
- These patterns were dubbed **periodic** patterns.
- Using these patterns he rearranged the table to put properties in alignment, hence **periodic table.**
- Blank spaces were left to make room for undiscovered elements.
- Naturally, we discovered and filled those in!

Moseley Cleans Up

- Henry Moseley, British Physicist
- Used early X-Ray diffraction techniques to refine our understanding on atomic structure and atomic numbers.
- He actually demonstrated that atomic numbers were not just assigned, but had a link to X-ray spectra! (Moseley's Law)
- Using this new information he re-arranged Mendeleev's Periodic Table into a more modern format, **by number** instead of weight.
- Naturally his new organization left gaps were filled in by future chemists.

Spectra Lines

- What relationships can you see when you look at your periodic table and these X-ray plates? Is there a connection?



Exercise

- Make a timeline of events including the publication of atomic theory, standardization of weights, and the development of the periodic table.

- Draw a sketch of the periodic table, and fill in the trend lines we learned about.

- How do all these trends interact? How are they connected? Draw an excellent concept map.

Arrangement

- Groups – Vertical up and down the table.

- Period – Horizontal, from right to left across the table.

Major Families

- Elements on the Periodic Table are arranged in **Families.**
- **Families** are grouped together because they have similar properties and characteristics.
- Why are they similar? Why do they share properties?
- Does it have to do with their configurations?

Metals

- Mostly silver/gray colors, with the exception of Cu and Au.
- Mostly solid, except for Hg, which is a liquid.
- High melting points, generally above 800 degrees C.
- Excellent conductors of heat and electricity.
- Misc. Physical – Very shiny/lustrous when polished, highly malleable and ductile.
- Reactive with Acids.
- Found in Groups 1-12, some in 13-16 under the step line

Non-Metals

- Highly variable colors.
- Solids and gases, except for Br.
- Low melting points.
- Poor conductors of heat and electricity.
- Very dull, brittle, often powdery.
- Not reactive with Acids.
- Found in Group 18, some in 14-17 above the step line.

Metalloids

- Silvery gray to black.
- Solids at room temperature.
- Variable melting points.
- Not good conductors of heat and electricity by themselves.
- Toes the line between metal/nonmetal.
- Reactions with acids vary.
- Found along the step line in Groups 13-17.

Alkali Metals

- Usually gray-white.
- Solids at room temperature.
- Low melting points.
- Not good conductors of heat and electricity.
- Considered metals.
- Extremely reactive! Never found freeform in nature, but is always in important compounds.
- Found in Group 1.

Alkaline Earth Metals

- Usually gray-white.
- Solids at room temperature.
- Low melting points.
- Not good conductors of heat and electricity.
- Considered metals.
- Almost reactive as the alkali metals. Harder and denser and stronger than the alkalis too. These are found throughout the Earth's crust.
- Found in Group 2.

Transition Metals

- Silvery Gray.
- Solids at room temperature, except Hg.
- High melting points.
- Good conductors of heat and electricity.
- Considered metals, obviously.
- High luster, very dense and very strong.
- Less reactive than the alkali/alkaline metals.
- Found in Group 3-12.

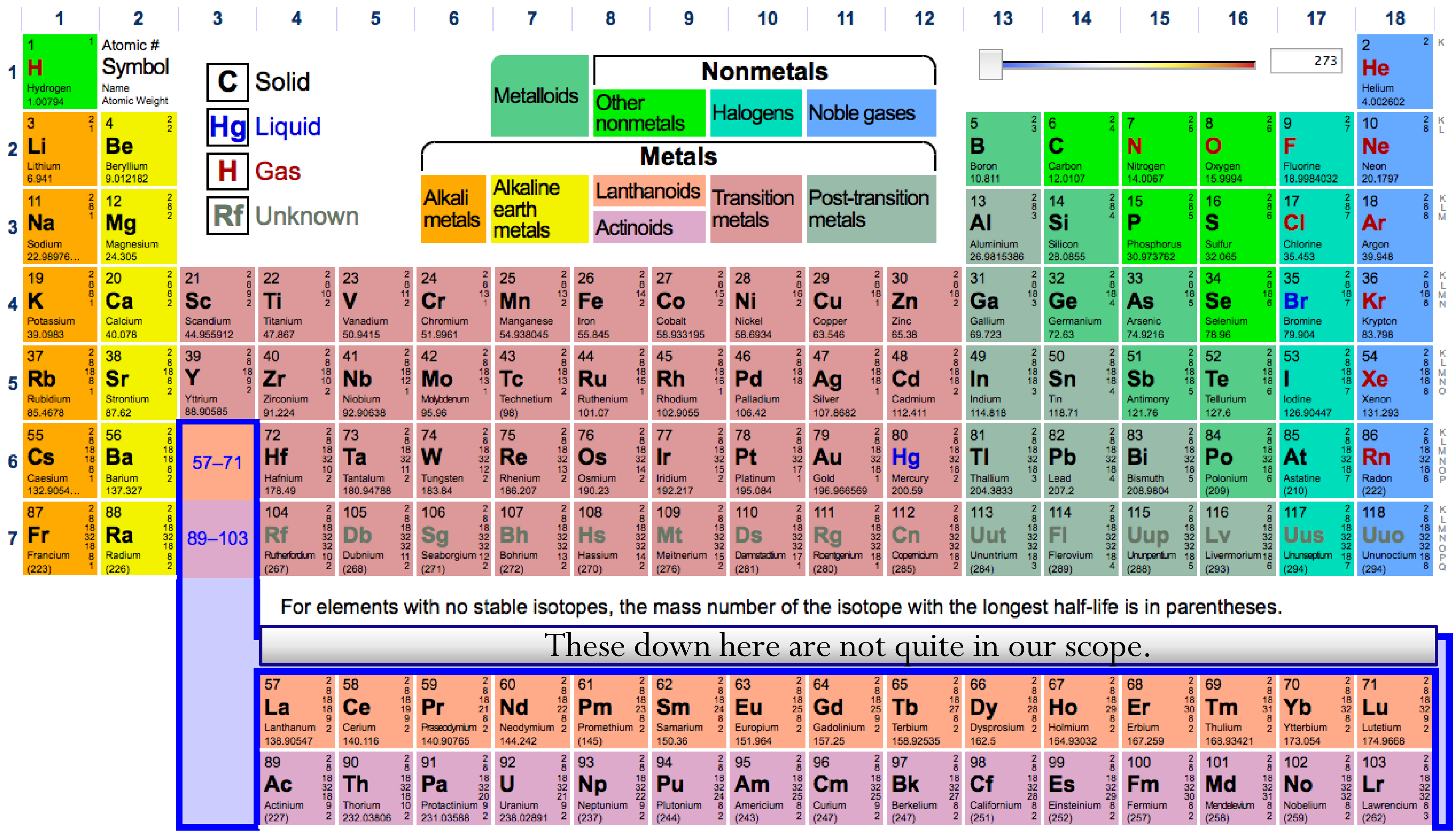
Halogens

- Various colors.
- Gases, except Br.
- Non-metals.
- Very reactive, often makes salts.
- Found in Group 17.

Noble Gases

- Gases.
- Light up in fun colors.
- Non-metals.
- Least reactive of all the elements!
- Found in Group 18.

The families



These down here are not quite in our scope.

Periodic Law

- When elements are arranged in order of increasing atomic number, there is a periodic pattern in their physical and chemical properties.
- There are numerous trends in the periodic table dealing with several different factors.
- You will need to know four of these trends and the reasons why they exist.

Periodic Trends

- Atomic Radius
  - How big is the atom? (Tricky Question)
- Ionic Radius
  - How big is an atom *when it is an ion?*
- Electronegativity
  - How strongly will this atom attract electrons?
- Ionization Energy
  - How much energy does it take to remove electrons?

Getting Real with Atoms

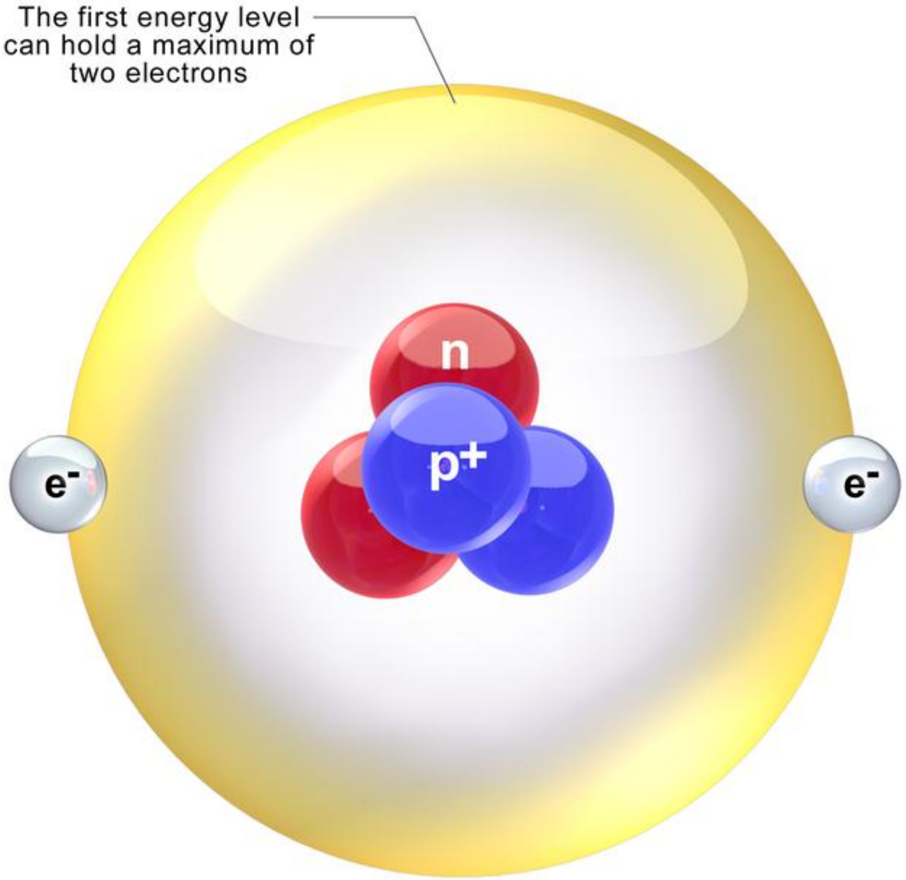
- What do atoms look like? What defines the boundaries?

- Atoms are not solid balls of matter. The line between atom and empty space is fuzzy.

- The nucleus is solid, but the electrons are in a cloud of potential places.

What do Atoms look like?

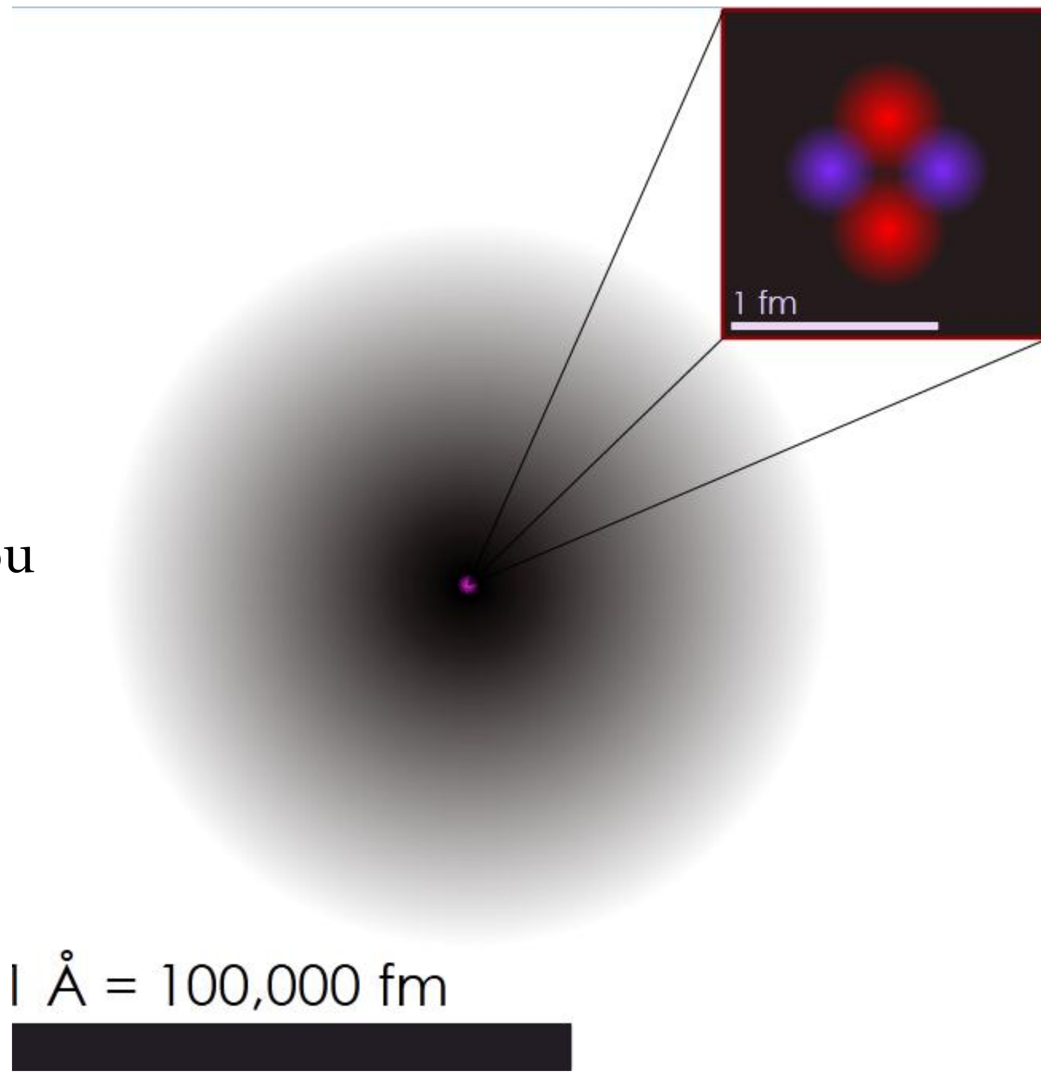
- We are used to seeing something like this;
- However, this is a simplified model designed to be easily read and printed on a page.
- Electrons do not have neat little orbits that can be easily measured and marked for us.



Helium, He
Atomic number: 2
Mass number: 4
(2 protons + 2 neutrons)
2 electrons

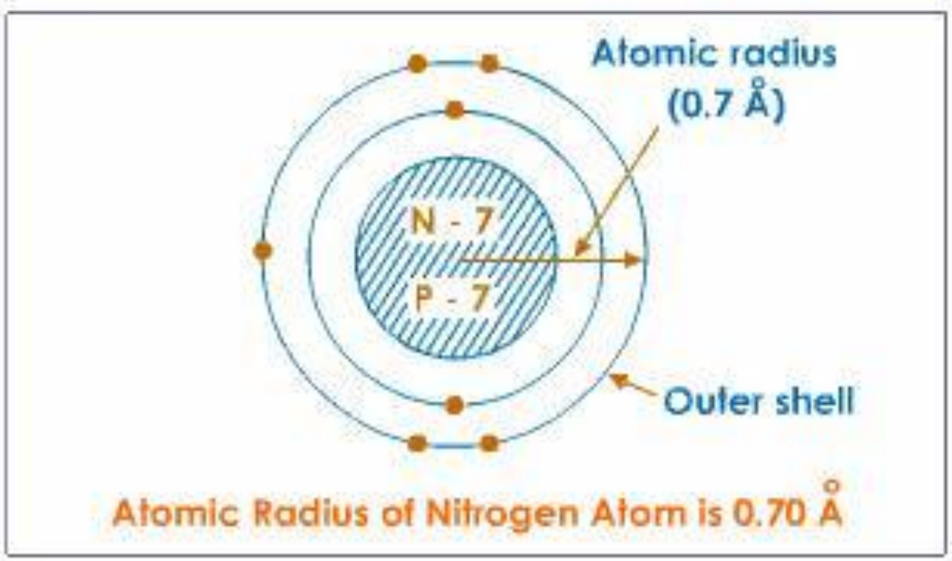
The Real Deal

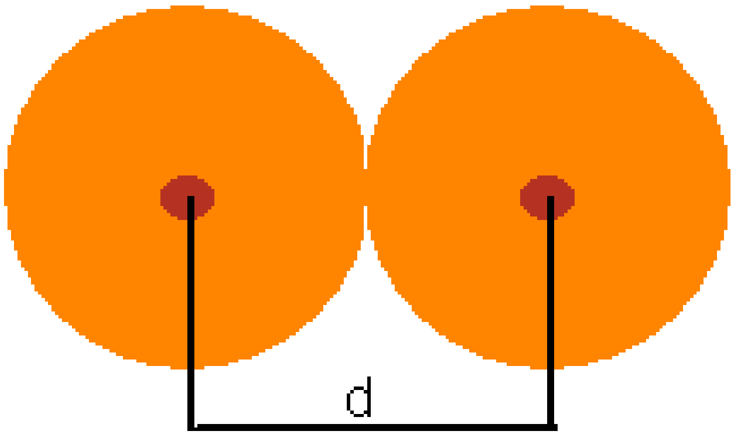
- This is a computer drawing made from data gathered about an atom of helium.
- That tiny dot in the center is the nucleus.
- The cloud is all the places you would detect those two electrons – the darker, the more likely they are to be there.
- The boundaries of an atom are not so clear-cut!



Atomic Radii

- How do we define it? Two options;
  - The distance from the atomic nucleus to the outermost stable electron shell in a neutrally charged atom
  - Half the distance between nuclei of atoms of the same element that are covalently bonded





$$r = d/2$$

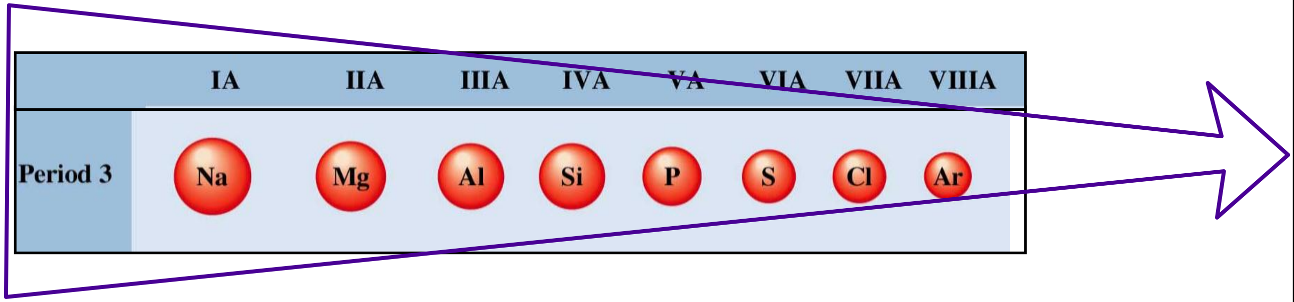
So.....those trends?

- Hold on – two important concepts are key here!

- First, **Nucleus Strength –** as the number of protons and neutrons in the nucleus increases, they can attract and hold on to electrons more strongly!
  - Bigger nucleus, more pull!
- Second, **Shielding –** electrons work together to 'block' or disperse the pull of the nucleus, the more electrons, the more they can weaken the pull of the nucleus.
  - More electrons, less pull!

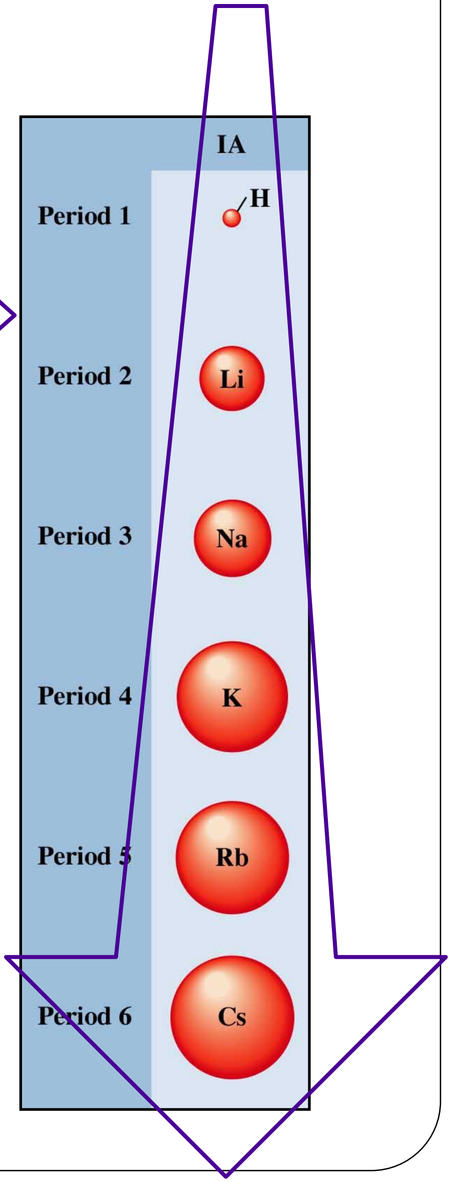
Atomic Radius Trends

- Across **periods,** the atomic radius **<u>decreases!</u>**



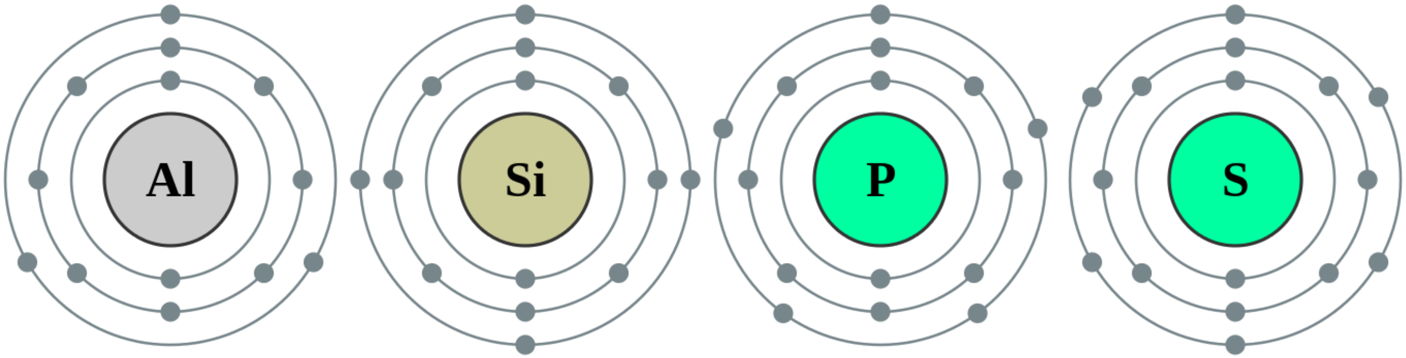
- Down **groups**, the atomic radius will **<u>increase</u>**.

- Why? It has to do with the interactions of the nucleus and the electrons!



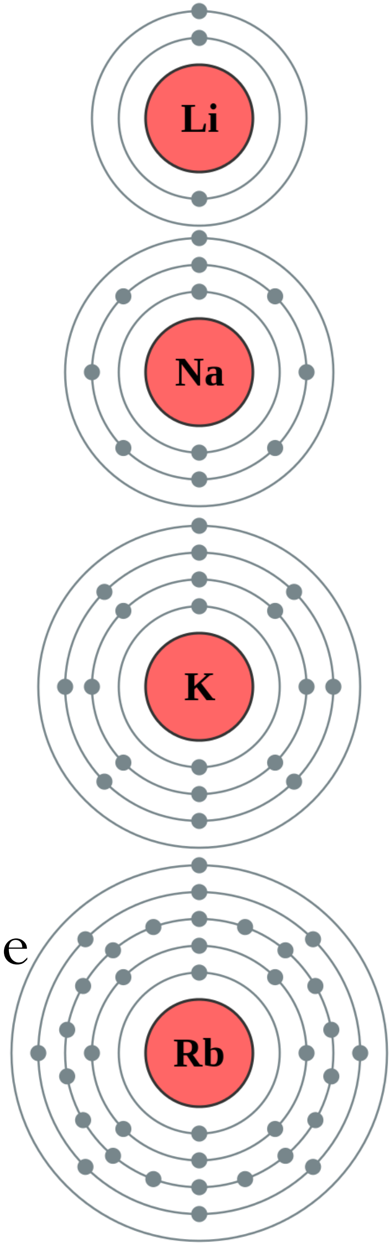
Atomic Radius Explained

- Across a **period** each element is gaining more electrons and protons. Where are they going?
  - Protons (and neutrons) are filling the nucleus up.
  - Electrons are filling up the electron shells.
- However, the electrons are not filling up a **new**, more distant shell!
- So the denser nucleus has a pull that increases faster than the electron's push – with a stronger pull, the electrons are forced closer!

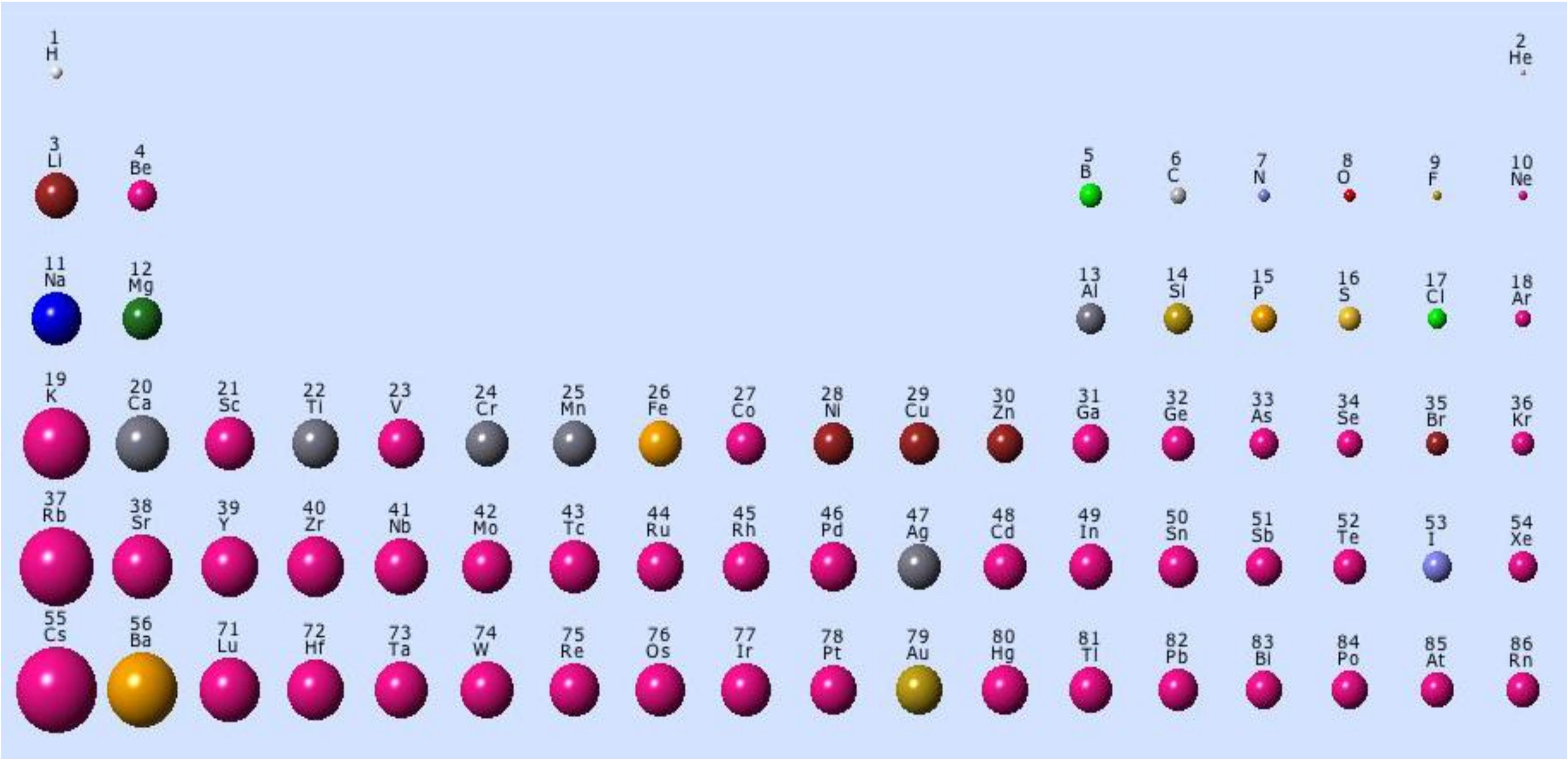


Atomic Radius Explained

- Down a **group** each element is gaining more electrons and protons. Where are they going?
  - Protons (and neutrons) are filling the nucleus up.
  - Electrons are filling up the electron shells.
- This time, the electrons are forced into **new**, more distant shells.
- While the nucleus is becoming more dense, the electrons are forced further and further away because the inner electron shells 'block' and 'push them out- they <u>cannot</u> get closer.



Atomic Radius Illustrated



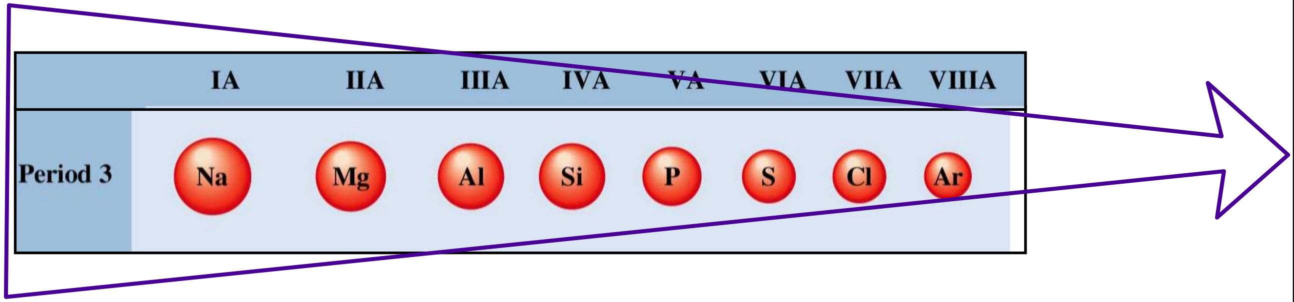
Ionic Radius Explained

- Ions are the results of losing or gaining electrons. Gains/losses of electrons change the balance of **push/pull.**
- When an atom gains electrons and becomes **negatively** charged, the other electrons get a stronger 'blocking' bonus and push further from the nucleus.
- When an atom loses electrons and becomes **positively** charged, the other electrons lose their 'blocking' bonus and get pulled closer.

Cations	**A**nions
Cramped	**A**part

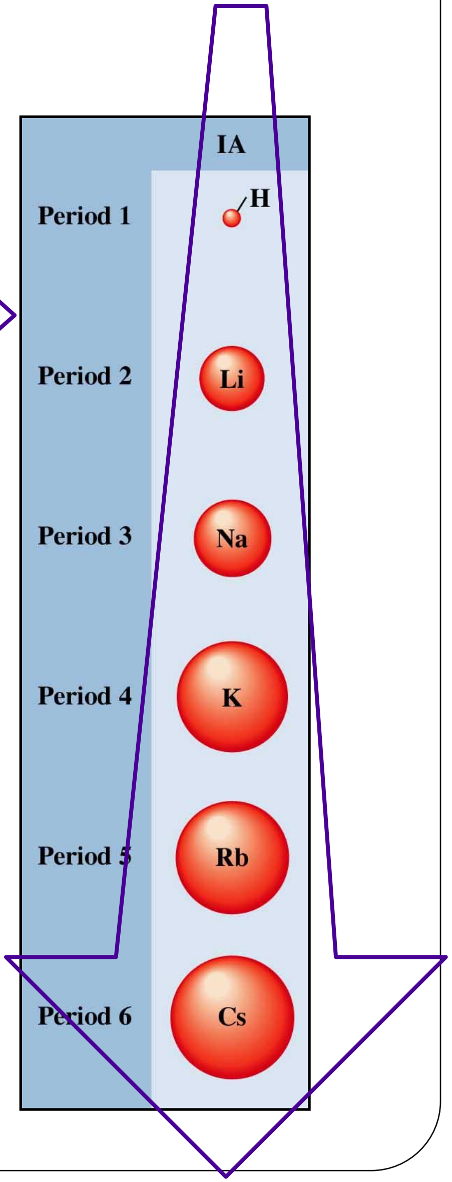
Ionic Radius Trends

- Across **periods,** the ionic radius **decreases!**



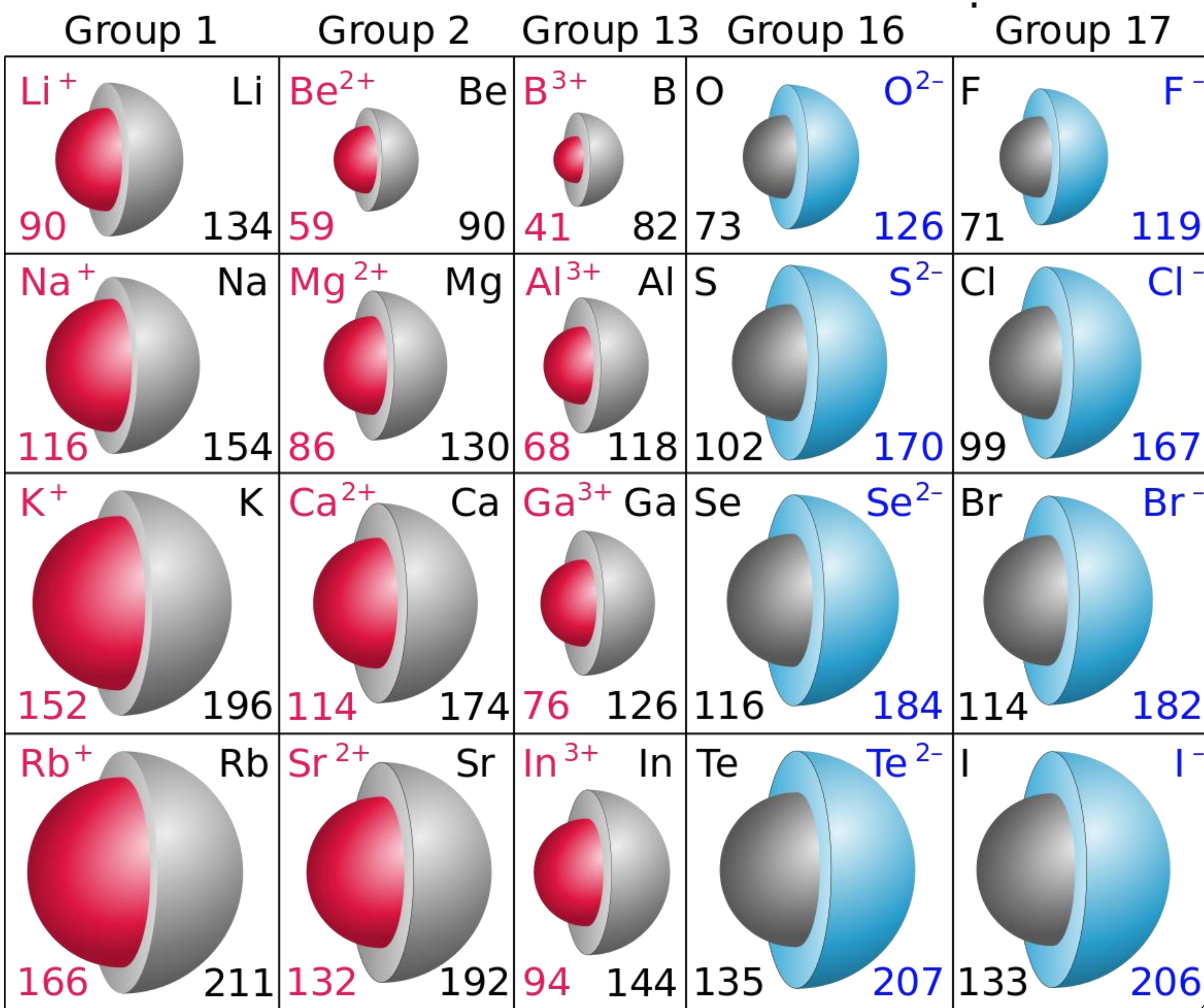
- Down **groups**, the ionic radius will **increase**.

- Ionic radius follows the same patterns as atomic radius for essentially the same reasons.



Ionic Radius Illustrated

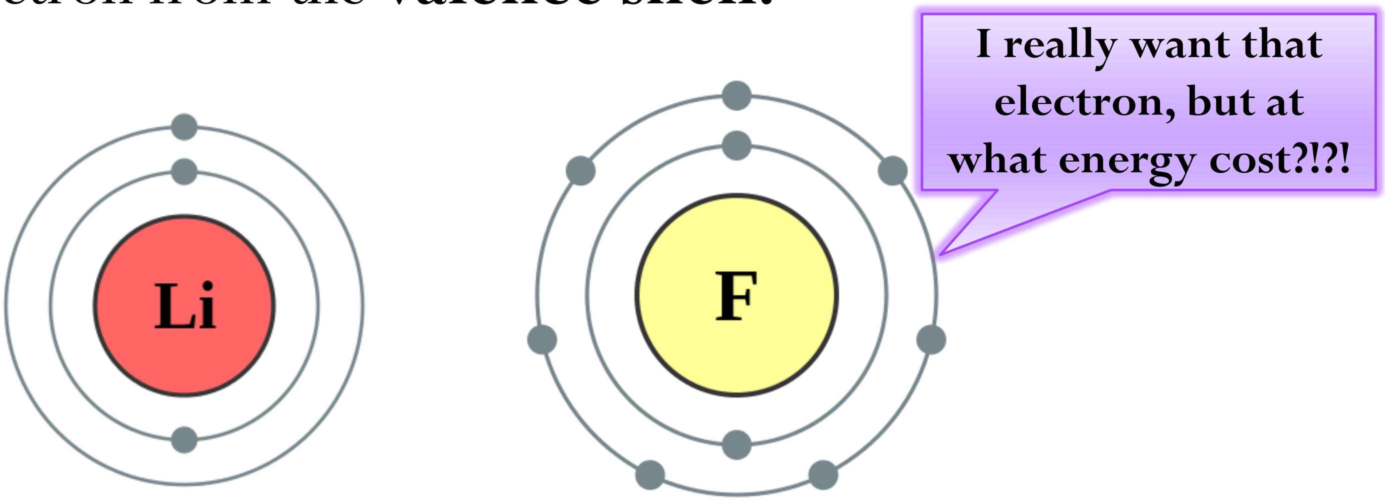
- Neutral Atomic Radius in Gray
- Cations in Red
- Anions in Blue



Ionization Energy

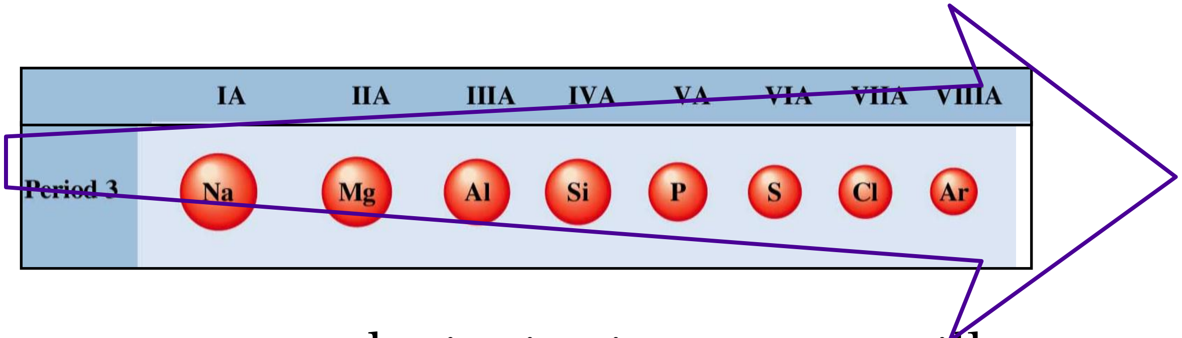
- The outermost shell of the atom is most important, it is the **valence shell** and is responsible for chemical reactivity.
- Every atom wants to have a complete **valence shell.**

- **Ionization Energy** is the energy needed to yank an electron from the **valence shell.**



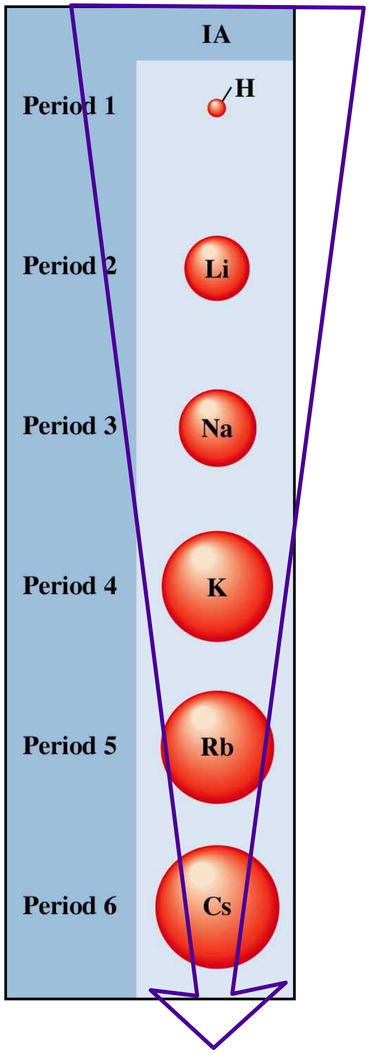
Ionization Energy Trends

- Across **periods,** the ionization energy **increases!**



- Down **groups**, the ionization energy will **decrease**.

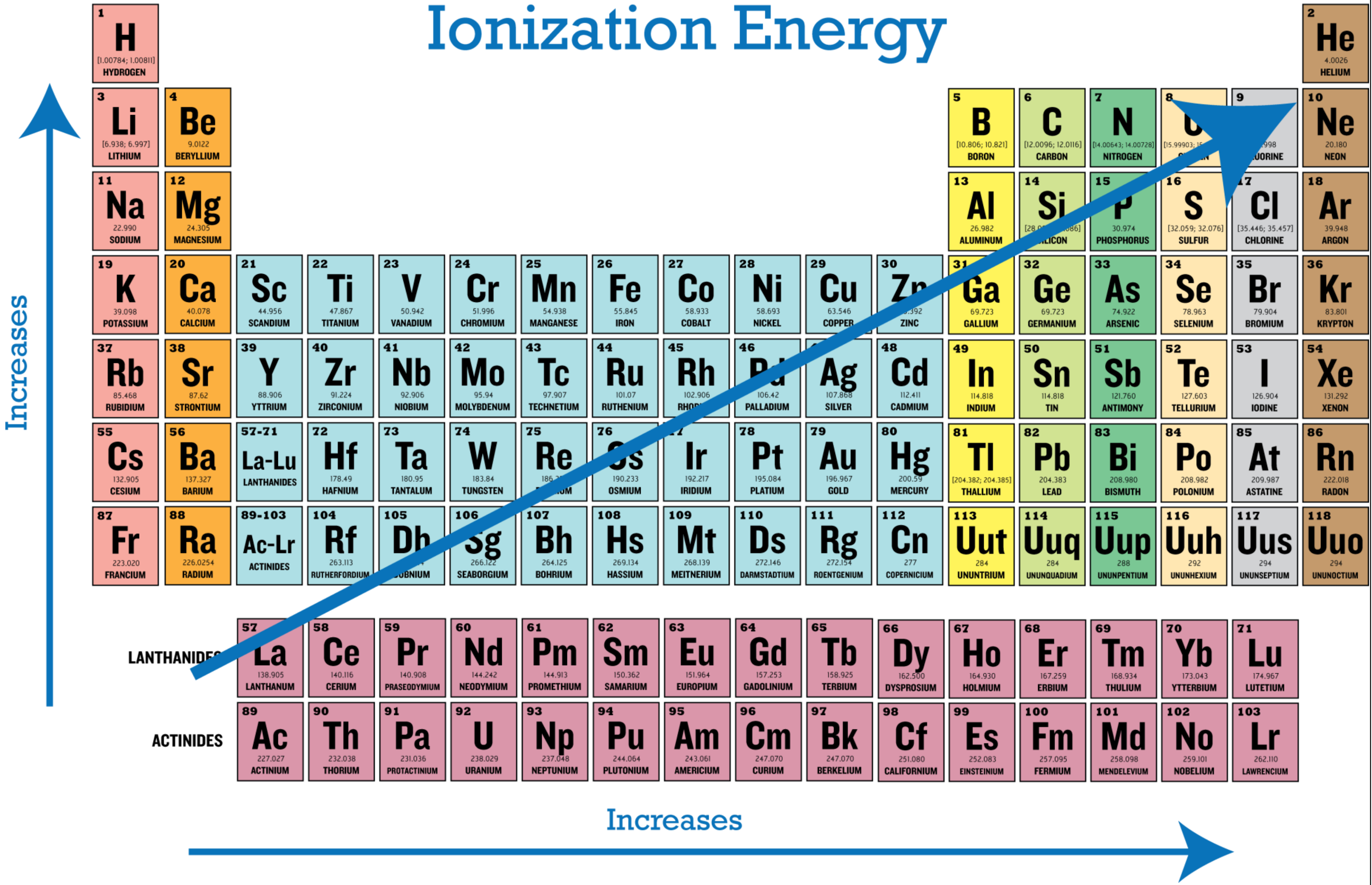
- Why? It has to do with the interactions of the nucleus and the electrons! (Again.)



Ionization Energy Explained

- Across a **period** the atomic radius is decreasing.
  - That means the **valence shell** is kept closer and closer to the nucleus.
- The closer the **valence electrons** are to the nucleus, the tighter it can hold onto those electrons!
- Down a **group** the atomic radius is increasing.
  - That means the **valence shell** is forced further and further away from the nucleus.
- The further the **valence electrons** from to the nucleus, the harder it is for them to be held.

Ionization Energy

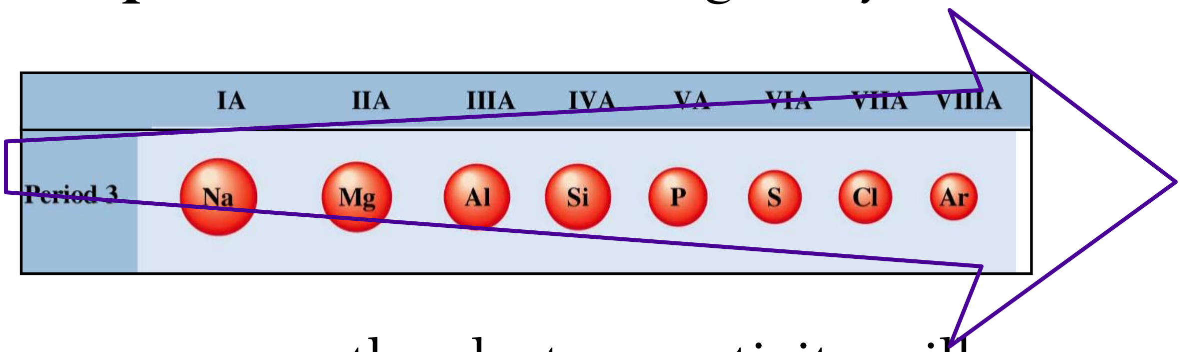


Electronegativity

- How attractive is the atom to electrons?
- What kind of force does the atom exert on other atoms as it tries to yank their electrons?

Electronegativity Trends

- Across **periods,** the electronegativity **<u>increases!</u>**



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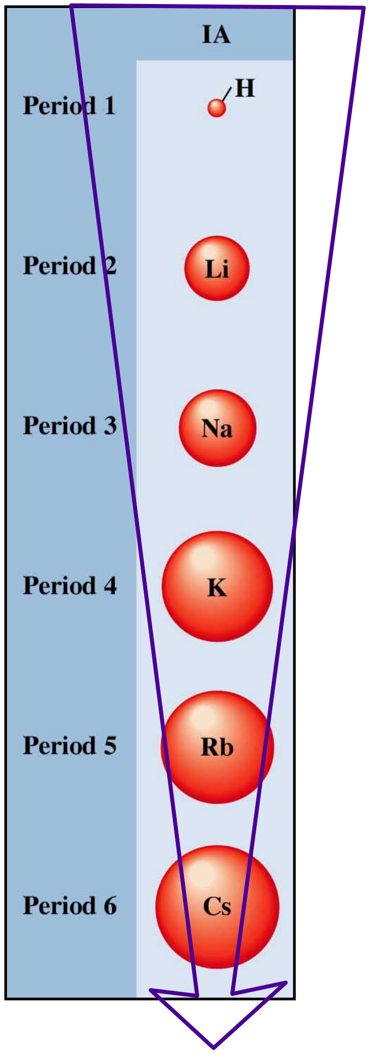

Electronegativity Explained

- Across a **period** the atomic radius is decreasing.
  - The electron shells are tight and close, exposing more of the nucleus' pulling field or zone of influence to the **target electron**

- The closer the target **electrons** can get to the nucleus, the harder it can pull on them.

- Down a **group** the atomic radius is increasing.
  - That means the **valence shell** is further away and the nucleus is already swarming with electrons – how hard can a nucleus pull when it is already surrounded?

- The further the **target electrons** are from the nucleus, the less force the nucleus can exert on it.

Electronegativity Illustrated

H 2.20																	He
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		